

VISIT...

LANZAROTE
Caliente.COM

SILVER IODIDE

[7783-96-2]

Formula AgI; MW 234.77

Uses

Silver iodide is used in cloud seeding for artificial rain making and in photography. Its colloidal suspension is used as a local antiseptic.

Physical Properties

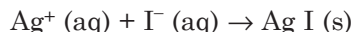
Light yellow hexagonal crystals or powder; darkens on exposure to light; density 5.68 g/cm³; melts at 558°C; vaporizes at 1,506°C; insoluble in water, most acids and ammonium carbonate solution; moderately soluble in concentrated solutions of alkali chloride, bromide, and thiosulfate; readily soluble in solutions of alkali cyanides, iodides and in hot concentrated hydriodic acid.

Thermochemical Properties

ΔH_f°	-14.8 kcal/mol
ΔG_f°	-15.8 kcal/mol
S°	27.6 cal/deg mol
C_p	13.6 cal/deg mol
ΔH_{fus}	2.25 kcal/mol
ΔH_{vap}	34.4 kcal/mol

Preparation

Silver iodide is prepared by adding a solution of sodium or potassium iodide to a hot solution of silver nitrate:



The precipitate is washed with boiling water. The preparation is done in the dark under ruby red light.

Analysis

Elemental composition: Ag 45.95%, I 54.05%. The salt is dissolved in hot concentrated nitric acid, diluted appropriately with water and analyzed for silver.

SILVER NITRATE

{7761-88-8}

Formula AgNO₃; MW 169.87; Synonym: lunar caustic

Uses

Silver nitrate is probably the most important silver salt. It is used to make most silver salts. It is used in photographic film, indelible ink, and hair dyeing. Other uses are in making silver mirrors, etching ivory, and as a catalyst

in preparing ethylene oxide. Silver nitrate is a titrant in all argentometric titration (Mohr titrations). In medicine, it is a topical anti-infective, an anti-septic, and its dilute solution is an eye lotion.

Physical Properties

Colorless, transparent, large rhombohedral crystals, or white small crystals; bitter, caustic metallic taste; odorless; pure compound is not sensitive to light but trace organics promote photo reduction, turning the salt to grayish black on exposure to light; density 4.35 g/cm³; melts at 212°C; decomposes at 440°C; very soluble in water, soluble in ethanol and acetone.

Thermochemical Properties

ΔH_f°	-29.7 kcal/mol
ΔG_f°	-7.98 kcal/mol
S°	33.7 cal/deg mol
C_p	22.3 cal/deg mol
ΔH_{fus}	2.75 kcal/mol

Preparation

Silver nitrate is prepared by dissolving silver metal in dilute nitric acid. The solution is evaporated and residue is heated to dull red heat with concentrated nitric acid to decompose impurities such as copper nitrate. Residue then is dissolved in water, filtered, and recrystallized to obtain pure silver nitrate.

Analysis

Elemental composition: Ag 63.50%, N 8.25%, O 28.25%. The salt is dissolved in water, diluted, and analyzed for silver. The nitrate ion, NO₃⁻, can be analyzed in aqueous solution by nitrate ion-selective electrode, ion chromatography, or colorimetry after reduction to NO₂⁻ ion with cadmium. The nitrate content of the salt is 36.50%.

Toxicity

Silver nitrate is toxic by all routes of exposure. Ingestion can cause severe gastroenteritis. Also, it is a severe irritant to eyes and skin.

LD₅₀ oral (rabbit): 800 mg/kg

SILVER(I) OXIDE

[20667-12-3]

Formula Ag₂O; MW 231.74

Synonyms: silver oxide; argentous oxide

Uses

Silver(I) oxide is used for polishing and coloring glass yellow. Also, it is used in purifying drinking water; as a catalyst; and as a germicide and parasiticide.

Physical Properties

Brownish-black cubic crystals; density 7.14 g/cm³ at 16°C; begins to decompose around 200°C, decomposition becoming rapid at 250 to 300°C; insoluble in water and ethanol; soluble in acids and alkalis; sparingly soluble in solutions of caustic alkalis; insoluble in alcohol.

Thermochemical Properties

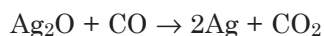
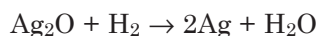
ΔH_f°	-7.43 kcal/mol
ΔG_f°	-2.68 kcal/mol
S°	29.0 cal/deg mol
C_p	15.75 cal/deg mol

Preparation

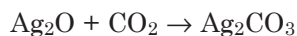
Silver(I) oxide is precipitated by mixing solutions of silver nitrate and caustic soda:

**Reactions**

When heated with hydrogen, carbon, carbon monoxide, or most metals silver(I) oxide is reduced to metallic silver:



Silver(I) oxide absorbs carbon dioxide in the presence of moisture producing silver carbonate:



The oxide dissolves in acids. Evaporation forms the silver salt.

Analysis

Elemental composition: Ag 93.10%, O 6.90%. The oxide is dissolved in nitric acid, diluted, and analyzed for silver. Its oxygen content may be measured by gravimetry following its reduction with hydrogen.

SILVER(II) OXIDE

[1301-96-8]

Formula AgO; MW 123.87

Synonyms: silver peroxide; argentic oxide; silver suboxide; Divasil

Uses

Silver(II) oxide is used to make silver oxide-zinc alkali batteries. Also, it is an oxidizing agent.

Physical Properties

Gray monoclinic or cubic crystals or powder; diamagnetic; semiconductor; density 7.48 g/cm³; decomposes to its elements above 100°C; insoluble in water (solubility 27 mg/L at 25°C); soluble in alkalis; decomposes in ammonia solution evolving nitrogen; dissolves in dilute acids with decomposition evolving oxygen; forms a brown solution in concentrated nitric acid, and forms intense green coloration in concentrated sulfuric acid.

Preparation

Silver(II) oxide is prepared by reacting silver nitrate with potassium persulfate in the presence of a base.

Analysis

Elemental composition: Ag 87.08%, O 12.92%. When dissolved in dilute nitric acid, oxygen is liberated immediately, which can be measured by GC or GC/MS (m/z 32). Acid solution may be analyzed for silver by AA, ICP, or other methods. When treated with ammonia solution, nitrogen is evolved which can be measured by GC or GC/MS (m/z 28).

SILVER SULFATE

[10294-26-5]

Formula Ag₂SO₄; MW 311.80

Uses

Silver sulfate is used as a catalyst to oxidize long chain aliphatic hydrocarbons in the determination of chemical oxygen demand (COD).

Physical Properties

Colorless crystals or powder; slowly darkens when exposed to light; density 5.45 g/cm³; melts at 652°C; decomposes at 1,085°C; slightly soluble in water; dissolves in nitric acid, concentrated sulfuric acid and ammonia solution.

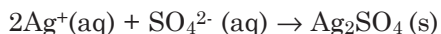
Thermochemical Properties

ΔH_f°	-171.1 kcal/mol
ΔG_f°	-147.8 kcal/mol
S°	47.9 cal/deg mol

Preparation

Silver sulfate is precipitated by adding sulfuric acid to a solution of silver

nitrate:



The precipitate is washed with hot water and preparation is under ruby red illumination.

Analysis

Elemental composition: Ag 69.19%, S 10.28%, O 20.52%. The salt is dissolved in nitric acid, the solution diluted, and analyzed for silver. It is very slightly soluble in water. The supernatant solution containing trace sulfate anion may be measured by ion chromatography or by treating with barium chloride followed by colorimetric measurement at 420 nm.

SILVER SULFIDE

[21548-73-2]

Formula Ag_2S ; MW 247.80

Synonym: argentous sulfide

Occurrence and Uses

Silver sulfide occurs in nature as mineral argentite. It is used in ceramics.

Physical Properties

Grayish-black orthogonal crystals or powder; density 7.23 g/cm³; Moh's hardness 2.3; melts at 825°C; insoluble in water; soluble in nitric and sulfuric acids.

Thermochemical Properties

ΔH_f°	-7.79 kcal/mol
ΔG_f°	-9.73 kcal/mol
S°	34.4 cal/deg mol
C_p	18.3 cal/deg mol
ΔH_{fus}	3.37 kcal/mol

Preparation

Mineral argentite is mined from mineral deposits, crushed, ground, and washed for use. In the laboratory, silver sulfide is obtained by passing hydrogen sulfide gas through a solution of silver nitrate. The precipitate is washed with hot water.

Analysis

Elemental composition: Ag 87.06% and S 12.94%. Silver sulfide is dissolved in nitric acid, the solution diluted and analyzed for silver. Also, it may be characterized nondestructively by x-ray diffraction.

SODIUM

[7440-23-5]

Symbol Na; atomic number 11; atomic weight 22.9898; a Group 1A (Group 1) alkali metal element; electron configuration $[\text{Ne}]3s^1$; valence +1; atomic radius 1.85Å; ionic radius, Na^+ in crystals 1.02Å (for a coordination number 6); ionization potential 5.139 eV; standard electrode potential, $E^\circ(\text{Na}^+ + e^- \leftrightarrow \text{Na})$ -2.71 V; one naturally-occurring stable isotope, Na-23 (100%); sixteen artificial radioactive isotopes in the mass range 19–22, 24–35; longest-lived radioisotope, Na-22, $t_{1/2}$ 2.605 year; shortest-lived isotope Na-35, $t_{1/2}$ 1.5 ms.

History, Occurrence, and Uses

Sodium was first isolated by Sir Humphry Davy in 1807 by electrolysis of caustic soda. In the following year, Gay Lussac and Thenard obtained metallic sodium by chemical reduction of caustic soda with iron at elevated temperatures. Deville, in 1854, prepared the metal by reduction of sodium carbonate and lime with charcoal at a temperature above the boiling point of sodium. Castner, in 1886, improved the chemical reduction process preparing the metal by heating sodium hydroxide with iron carbide at high temperature. Five years later he patented a process based on electrolytic reduction of sodium hydroxide. The first major commercial plant was set up in 1921 with the introduction of Downs cell.

The element derived its name from the Latin word *sodanum* meaning “headache remedy.” Its symbol Na was derived from the Latin word, *natrium*.

Sodium is the sixth most abundant element on earth. It comprises about 2.6% weight of the earth’s crust. Its salt, sodium chloride, is the major component of seawater. The concentration of sodium in seawater is 1.08%. As a very reactive element, sodium is never found in free elemental form. It occurs in nature in many minerals such as cryolite, amphibole, zeolite, sodalite, and soda niter. Sodium chloride (NaCl) is the most common salt of sodium. Some other important salts are caustic soda (NaOH), soda ash (Na_2CO_3), baking soda (NaHCO_3), Chile saltpeter (NaNO_3), borax ($\text{Na}_2\text{B}_4\text{O}_7 \cdot 10\text{H}_2\text{O}$), sodium thiosulfate ($\text{Na}_2\text{S}_2\text{O}_3$), sodium sulfate (Na_2SO_4), and sodium phosphates.

Metallic sodium is a strong reducing agent, used in many organic syntheses. It is used in the manufacture of sodamide, sodium peroxide, and esters. Other uses are in purifying molten metals, to descale metal, to improve structure of certain alloys, and as a heat transfer agent, for example, in nuclear reactors. Sodium is useful in producing other metals, such as titanium. It is used in sodium vapor lamps in small amounts. Sodium wire is used to remove traces of water from organic solvents.

Physical Properties

Soft, bright, silvery metal; malleable, can be readily cut with a knife or extruded as wire; liquid sodium in inert atmosphere appears like mercury; blue vapor, appears brilliant green at high temperatures; imparts golden-yellow color to flame; body-centered cubic structure; paramagnetic; density 0.97

g/cm³; melts at 97.72°C; vaporizes at 883°C; vapor pressure 1 torr at 439°C and 5 torr at 511°C; electrical resistivity 4.69 microhm-cm at 20°C and 6.60 microhm-cm at its melting point; viscosity 0.680 centipoise at 100°C; surface tension 192 dyne/cm at its melting point; neutron absorption cross section 0.505 barns; reacts violently with water; soluble in liquid ammonia forming a deep blue solution; soluble in ethylenediamine.

Thermochemical Properties

ΔH_f° (cry)	0.0 kcal/mol
ΔH_f° (gas)	25.7 kcal/mol
ΔG_f° (gas)	18.4 kcal/mol
S° (cry)	12.3 cal/deg mol
S° (gas)	36.7 cal/deg mol
C_r (cry)	6.74 cal/deg mol
C_r (gas)	4.97 cal/deg mol
ΔH_{fus}	0.62 kcal/mol
Thermal conductivity (at 27°C)	1.41 W/cm K
Coefficient of linear expansion (at 25°C)	$71 \times 10^{-6}/^\circ\text{C}$

Production

Sodium metal is produced by both electrolytic and chemical reduction processes. All commercial processes employed today are based on electrolytic methods. Such processes are in wide use since Davy prepared the metal the first time in 1807.

There are two electrolytic methods that are of major importance. One involves the electrolysis of fused sodium chloride using the Downs cell. This method currently is most prevalent. The Downs cell consists of a steel cell with brick lining containing the fused bath. The multiple electrode arrangement consists of four cylindrical graphite anodes that project upward from the base of the cell. Each anode is surrounded by a diaphragm of iron gauge and a steel cathode.

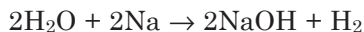
Fused sodium chloride is electrolyzed at bath temperature varying between 565 to 600°C at a cell voltage of 5.7 to 7 V and the cell current varying from 25 to 35 kA. The cathode current density is mostly about 9.8 kA/m². Often calcium chloride is added to sodium chloride in the cell bath to lower its melting point. Calcium is largely removed from sodium by filtration at about 110°C. Other electrolyte compositions have been used in which calcium is partially or fully replaced. The cell feed must be free of sulfate and other impurities.

Electrolysis of fused sodium hydroxide has been achieved successfully with a Castner cell. The Castner cell was used in commercial production prior to introduction of Downs cell. The cell is operated at a bath temperature $320 \pm 10^\circ\text{C}$, at 9.0 ± 0.5 amp current and a voltage of 4.3 to 5.0 V. The cathode current density is about 10.9 kA/m². The cell consists of a copper cathode and a nickel anode and a cylindrical iron-gauge diaphragm placed between the electrodes. The cell reactions are as follows:





Water generated at the anode diffuses through the diaphragm and goes to the cathode, reacting with sodium to form sodium hydroxide.

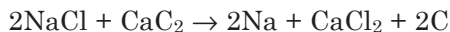
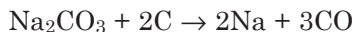


The overall change may be represented as:



Because water is reacting with sodium produced at the cathode, the yield of sodium is reduced almost by 50%. Lesser yield is the major disadvantage of the Castner process. At present, this process is not used commercially.

Thermal reduction processes are not being practiced anywhere in the world at present for large-scale production of sodium. Such methods, however, can be conveniently adapted for laboratory preparation of metallic sodium. Sodium can be prepared by thermal reduction of its hydroxide, carbonate, or chloride at elevated temperatures. These salts are heated with carbon, calcium carbide, iron carbide, ferrosilicon, or other reducing agents at temperatures above 800°C under vacuum:



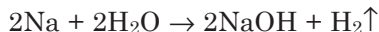
Reactions

Sodium is a highly reactive metal. Most reactions are violent. Sodium ignites in air when heated at 120°C, burning with a yellow flame, forming a dense white smoke with an acrid odor. It forms three oxides, the monoxide, Na_2O ; the peroxide, Na_2O_2 ; and the superoxide, NaO_2 . When heated below 160°C under a limited supply of oxygen, sodium monoxide, Na_2O , is the major product. At 250 to 300°C in adequate oxygen, sodium forms its peroxide, Na_2O_2 , along with trace amounts of superoxide, NaO_2 . When heated above 300°C under oxygen pressure, the metal forms the superoxide, NaO_2 . Sodium dissolved in liquid ammonia reacts with oxygen to produce the superoxide. The reaction is rapid, but the product is impure. Sodium also reacts with ozone forming an unstable ozonide, NaO_3 .

Sodium combines with hydrogen forming sodium hydride, NaH . The reaction is slow at ambient temperature but proceeds rapidly above 200°C when the metal is dispersed or spread over the surface of an inert solid (such as a hydrocarbon). Sodium and hydrogen react with aluminum powder to form sodium aluminum hydrides. Two such complex hydrides, the tetrahydride, NaAlH_4 , and the hexahydride, Na_3AlH_6 , are produced. The nature of the prod-

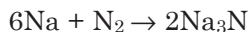
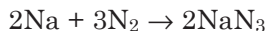
uct depends on reaction conditions. Under high hydrogen pressure and higher aluminum to sodium ratio, formation of the tetrahydride is favored. Such reactions are catalyzed by trialkyl aluminum.

Sodium reacts violently with water liberating hydrogen:

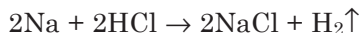


The reaction is highly exothermic; the heat of reaction is about 99 kcal/mol. In a closed system excessive pressure generated can cause an explosion. The reaction can be well controlled in inert atmosphere if the heat of reaction is dissipated. The reaction occurs vigorously even at -80°C .

Sodium is stable in nitrogen at most temperatures. Reaction, however, occurs at very high temperatures or when nitrogen is activated by electric discharge. The products are sodium azide, NaN_3 , and sodium nitride, Na_3N :



Reactions with dilute mineral acids can be vigorous to violent with liberation of hydrogen:



Sodium is a powerful reducing agent. It reduces a number of metal oxides to metals at high temperatures. Examples are oxides of iron, copper, zinc, cadmium, mercury, chromium, titanium, and many other metals. It does not reduce oxides of lithium, magnesium, or calcium.

Sodium reduces most metal chlorides to metals. Thus, when heated with titanium or zirconium tetrachloride, sodium converts the halides to free metals. Chlorides of calcium, magnesium, and potassium are only partially reduced.

Sodium dissolves in liquid ammonia forming an unstable blue solution. The reaction is slow. Sodium amide and hydrogen are generated:

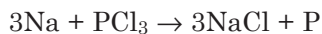


This reaction is catalyzed by iron, cobalt, and nickel. Rate of reaction depends on temperature and concentration of sodium in liquid ammonia. At a temperature of -41.6°C and high concentration, the solution separates into two liquid phases that consist of a deep blue dilute solution at the bottom that is low in sodium, and a lighter solution of metallic bronze color on the top with a high sodium. Molten sodium reacts with ammonia gas at 300 to 400°C to form sodium amide.

Sodium reacts with carbon monoxide at 250 to 340°C forming sodium carbonyl, $(\text{NaCO})_6$. At higher temperatures, sodium carbide Na_2C_2 is formed. With acetylene the products are sodium acetylide, $\text{NaC}\equiv\text{CH}$ and disodium

acetylide $\text{NaC}\equiv\text{CNa}$ (also known as sodium carbide). The latter compound also is obtained by heating sodium metal with sodium carbonate at 500 to 700°C.

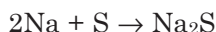
Sodium reacts with phosphorus on heating to form sodium phosphide, Na_3P . When ignited with phosphorus in the presence of air, sodium phosphate, Na_3PO_4 , is obtained. When heated with phosphorus trichloride, sodium reduces the latter compound to elemental phosphorus:



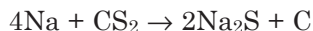
On the other hand, heating the metal with phosphorus pentachloride yields sodium phosphide:



Sodium combines with sulfur, selenium, and tellurium at high temperatures forming binary compounds. With sulfur the product is sodium sulfide:

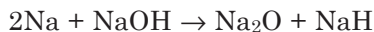


Sodium sulfide also is produced when the metal is heated with carbon disulfide. The reaction is violent:

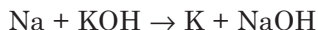


Sodium combines with all halogens forming sodium halides. The metal ignites with fluorine, forming hydrogen fluoride. Thin metal film reacts readily with chlorine and bromine at ordinary temperatures. Molten sodium burns in chlorine producing sodium chloride. The metal reacts with iodine, only in vapor phase, forming sodium iodide.

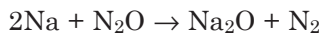
Sodium reacts with caustic soda at temperatures between 300 to 385°C:



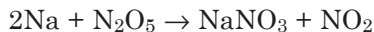
Reaction with caustic potash is complicated, involving several intermediates that finally yield potassium metal and sodium hydroxide:



Reaction with nitrous oxide yields sodium oxide:

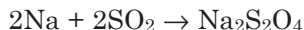


Reaction with liquid nitrogen pentoxide at low temperatures forms sodium nitrate and nitrogen dioxide:



Sodium reacts rapidly with hydrogen sulfide in the presence of moisture to form sodium sulfide. With dry hydrogen sulfide the reaction is slow.

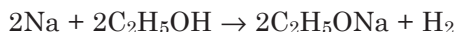
The molten metal reacts violently with sulfur dioxide to form sodium hydrosulfite, $\text{Na}_2\text{S}_2\text{O}_4$:



Sodium forms alloys with a number of metals including lead, chromium, mercury, aluminum, silicon, and iron. With mercury, it forms sodium amalgam. Sodium-lead alloy is commercially used to produce tetraethyllead, which was used historically as an additive to gasoline:

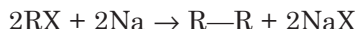


Sodium reacts with lower primary alcohols forming its alkoxide:

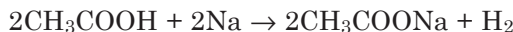


Reaction is slow with secondary and tertiary alcohol.

Sodium displaces halogens from alkyl halides forming alkanes (Wurtz reaction):



Carboxylic acids may react with sodium forming sodium salts, liberating hydrogen or they may decompose:



Sodium reacts with naphthalene in dimethyl ether to form a dark green reactive complex. This addition product, naphthalenesodium, $\text{C}_{10}\text{H}_8\text{Na}$, is stabilized by solvation with ether. Anthracene, phenanthrene, biphenyl, and many other aromatics form similar complexes with sodium in the presence of methylethyl ether, tetrahydrofuran, dioxane, and other ethers.

Analysis

All sodium compounds impart a golden yellow color to flame. Sodium can be identified spectroscopically by characteristic line spectra. Trace sodium may be measured quantitatively by flame atomic absorption or flame emission photometric method. The element may be measured at 589 nm using an air-acetylene flame. If using an ICP-atomic emission spectrophotometer, sodium may be measured at 589.00 or 589.59nm. Metallic sodium may be analyzed quantitatively by treating with ethanol and measuring the volume of hydrogen liberated.

Hazard

Sodium is a highly reactive metal. It ignites in air and reacts violently with

water. Many of its reactions are explosive (see Reactions). It should be stored under kerosene or hydrocarbon solvents. Contact with skin can cause serious burns. Contact with the eyes can cause blindness.

SODIUM ACETATE

[127-09-3]

Formula: CH_3COONa ; MW 82.035; also forms a stable trihydrate, $\text{CH}_3\text{COONa} \cdot 3\text{H}_2\text{O}$ [6131-90-4], MW 136.08

Uses

Sodium acetate is a mordant in dyeing. Other applications are in photography, as an additive to food, in purification of glucose, in preservation of meat, in tanning, and as a dehydrating agent. In analytical chemistry it is used to prepare buffer solution.

Physical Properties

Anhydrous salt is a colorless crystalline solid; density 1.528 g/cm³; melts at 324°C; very soluble in water; moderately soluble in ethanol.

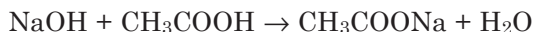
The colorless crystalline trihydrate has a density 1.45 g/cm³; decomposes at 58°C; is very soluble in water; pH of 0.1M aqueous solution is 8.9; moderately soluble in ethanol, 5.3 g/100mL.

Thermochemical Properties

ΔH_f°	-169.4 kcal/mol
ΔG_f°	-145.2 kcal/mol
S°	23.4 cal/deg mol
C_p	19.1 cal/deg mol

Preparation

Sodium acetate is prepared by reacting sodium hydroxide or sodium carbonate with acetic acid in aqueous solution. The solution is evaporated to obtain hydrated crystals of sodium acetate.



SODIUM AMIDE

[7782-92-5]

Formula NaNH_2 ; MW 39.013

Synonym: sodamide

Uses

Sodium amide is a dehydrating agent. It is used in preparing sodium cyanide and hydrazine, and in many organic synthetic reactions such as Claisen condensations, alkylations of ketones and nitriles, and in ammonolysis reactions.

Physical Properties

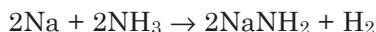
White crystalline powder with odor of ammonia; orthogonal crystals; density 1.39 g/cm³; melts at 210°C; begins to volatilize at 400°C; decomposes at 500°C; decomposed by water and hot alcohol; in fused state it dissolves zinc, magnesium and other metals, as well as, quartz, glass, and silicates.

Thermochemical Properties

ΔH_f°	-29.6 kcal/mol
ΔG_f°	-15.3 kcal/mol
S°	18.4 cal/deg mol
C_p	15.8 cal/deg mol

Preparation

Sodium amide is prepared by passing dry ammonia gas over sodium metal at 350°C:

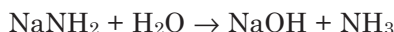


Also, it may be prepared by reacting sodium metal with liquid ammonia in the presence of a catalyst such as iron(III) nitrate. The compound must be stored in well-sealed containers free from air or moisture.

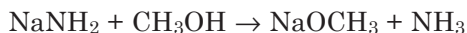
Reactions

Sodium amide dissociates to its elements, sodium, nitrogen, and hydrogen at temperatures between 500 and 600°C.

Its reaction with water is violent, forming sodium hydroxide and ammonia:



With alcohol the reaction is moderate forming sodium alkoxide and ammonia:

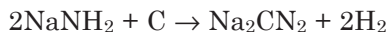


When heated with nitrous oxide at 200°C the products are sodium azide, caustic soda, and ammonia:



Sodium amide reacts with carbon at 800°C to form sodium cyanamide and

hydrogen:



Analysis

Elemental composition: Na 58.93%, N 35.90%, H 5.17%. The compound may be decomposed cautiously with water (reaction is violent) under cooling to yield sodium hydroxide and ammonia. (Or it may be decomposed with anhydrous alcohol to form ammonia and sodium alcoholate. The alcoholate then may be treated with water to form sodium hydroxide). Ammonia liberated is dissolved in water and the solution is measured using an ammonia-selective electrode. Alternatively, ammonia is collected over boric acid solution containing a small quantity of methyl red indicator. The solution is titrated with a standard solution of sulfuric acid. Sodium hydroxide is measured by titration with a standard solution of hydrochloric or sulfuric acid.

Hazard

Sodium amide is a flammable solid. It undergoes violent reactions with oxygen (air), water, and oxidants. Also, it explodes when heated, crushed or grinded. If not properly sealed, it can become explosive on storage, the warning sign for which is development of yellow or brownish color. Such material may be destroyed safely by covering with benzene or toluene and slowly adding ethanol while stirring.

The compound is a strong irritant to skin and eye.

SODIUM AZIDE

[26628-22-8]

Formula NaN_3 ; MW 65.01

Uses

Sodium azide is used to make lead azide and hydrazoic acid, and as a propellant for automotive safety bags. It also is used as an antihypertensive agent to control blood pressure.

Physical Properties

Colorless hexagonal crystals; density 1.846 g/cm^3 at 20°C ; decomposes on heating to produce sodium and nitrogen; also decomposes in vacuum; soluble in water partially converting to hydrazoic acid, solubility in water, 41.7 g/100mL ; slightly soluble in alcohol, 0.316 g/100mL at 16°C ; soluble in liquid ammonia.

Thermochemical Properties

ΔH_f° (cry)	5.19 kcal/mol
ΔG_f° (cry)	22.41 kcal/mol
S° (cry)	23.15 cal/deg mol

C_p (cry) 18.31 cal/deg mol

Preparation

Sodium azide is prepared by reacting sodium amide with nitrous oxide. The amide is heated with nitrous oxide at 200°C or its solution in liquid ammonia is treated with nitrous oxide at ambient temperature:



Analysis

Elemental composition: Na 35.36%, N 64.64%. The salt is dissolved in water, sufficiently diluted, and analyzed for sodium (see Sodium). The solid powder is decomposed cautiously and liberated nitrogen is measured by GC-TCD or GC/MS. The characteristic mass for N_2 is 28.

Hazard

Sodium azide is a toxic as well as an explosive substance (Patnaik, P. 1999. *A Comprehensive Guide to the Hazardous Properties of Chemical Substances*, 2nd ed. New York: John Wiley & Sons). Although inert to shock, violent decomposition can occur when heated at 275°C. Contact of solid or solution with lead and copper must be avoided. Reactions with halogens, carbon disulfide, or chromyl chloride can be explosive. Dissolution in water produces toxic vapors of hydrazoic acid. The salt is an acute poison causing headache, hypotension, hypothermia, and convulsion.

LD₅₀ oral (rats): 27 mg/kg

SODIUM BICARBONATE

[144-55-8]

Formula NaHCO_3 ; MW 84.007

Synonyms: baking soda; sodium hydrogen carbonate; sodium acid carbonate

Uses

Sodium bicarbonate is an ingredient of baking powder. It also is used in making effervescent salts and beverages, artificial mineral waters, and several other sodium salts. It is used in fire extinguishers, in gold plating, in cleaning formulations, in preventing mold growth on timber, in mouthwash, and as a laboratory reagent. In medicine it is used in antacids and alkalizers.

Physical Properties

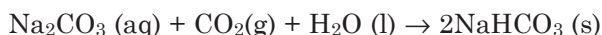
White crystalline powder or granules; monoclinic crystals; density 2.20 g/cm³; decomposes around 50°C, begins to lose carbon dioxide; converts to sodium carbonate at 100°C; soluble in water, 10g/100 mL at 20°C; slowly decomposes to CO_2 and Na_2CO_3 in aqueous solution at ambient temperature; decomposes to Na_2CO_3 in boiling water; aqueous solution slightly alkaline; pH of 0.1M solution at 25°C is about 8.3; insoluble in alcohol; decomposes in acids.

Thermochemical Properties

ΔH_f°	-227.2 kcal/mol
ΔG_f°	-203.4 kcal/mol
S°	24.3 cal/deg mol
C_p	20.9 cal/deg mol

Preparation

Sodium bicarbonate is prepared by passing carbon dioxide into a saturated solution of sodium carbonate. The bicarbonate, being less soluble than carbonate, precipitates:



Also, sodium bicarbonate is obtained as a by-product of sodium carbonate manufacture using the Solvay process (see Sodium Carbonate).

SODIUM CHLORIDE

[7647-14-5]

Formula NaCl; MW 58.443

Synonyms: common salt; salt; rock salt; halite; table salt.

Occurrence and Uses

Sodium chloride is widely distributed in nature. Oceans are the vast source of sodium chloride. It occurs in seawater at an average concentration of 2.68 wt%. It also occurs in many inland saline waters and in salt deposits in sedimentary rocks, as the mineral halite.

Sodium chloride is probably the most important salt of both sodium and chlorine. Sodium chloride, common table salt, is an essential component of most food preparation, imparting flavor to food and providing the sodium nutritional requirement. Also, it is used for preserving food. Therapeutically, NaCl solution is used to combat dehydration as an electrolyte replenisher, and it is an emetic.

The most important applications of sodium chloride in the chemical industry are in making a number of important industrial chemicals such as hydrochloric acid, sodium hydroxide, sodium carbonate, and metallic sodium. It is the starting material in manufacturing these substances. Other uses are in dyeing and printing fabrics, glazing pottery, in making soap, and for curing hides. Sodium chloride is a component of many freezing mixtures.

Physical Properties

White granular crystals or powder; large crystals are colorless, transparent, or translucent; saline taste; cubic structure; refractive index 1.5442; density 2.165 g/cm³; melts at 801°C; vaporizes at 1,413°C; soluble in water, 35.7g/100mL at 0°C and 39.1 g/100mL at 100°C; aqueous solution neutral; soluble in glycerol, ethylene glycol, and formic acid; sparingly soluble in

methanol (1.49 g/100 mL) and liquid ammonia (2.15 g/100mL); insoluble in hydrochloric acid.

Thermochemical Properties

ΔH_f° (cry)	−98.27 kcal/mol
ΔH_f° (gas)	−42.22 kcal/mol
ΔG_f° (cry)	−91.82 kcal/mol
ΔG_f° (gas)	−47.00 kcal/mol
S° (cry)	17.24 cal/deg mol
S° (gas)	54.90 cal/deg mol
C_p (cry)	12.07 cal/deg mol
C_p (gas)	8.55 cal/deg mol

Production

Sodium chloride is produced by solar evaporation of seawater or brine from underground salt deposits. It also is produced by mining rock salt. The commercial product contains small amounts of calcium and magnesium chlorides.

Analysis

Elemental composition: Na 39.34%, Cl 60.66%. Aqueous solution may be analyzed for sodium by various instrumental methods (see Sodium) and for chloride ion by ion chromatography or chloride-ion selective electrode. Alternatively, the chloride ion may be measured by titration with a standard solution of silver nitrate using potassium chromate as indicator. Also, the salt can be identified by its physical properties.

SODIUM BISULFATE

[7681-38-1]

Formula NaHSO_4 ; MW 120.06; forms a monohydrate, $\text{NaHSO}_4 \cdot \text{H}_2\text{O}$

Synonyms: sodium hydrogen sulfate; sodium acid sulfate; niter cake

Uses

Sodium bisulfate is used for pickling metals; bleaching leather; carbonizing wool; in carbonic acid baths, and manufacturing magnesia cements

Physical Properties

Colorless crystals; triclinic structure; density 2.435g/cm³ at 13°C; melts above 315°C; decomposes on further heating; soluble in water, 28.6 g/100mL at 25°C; highly soluble in boiling water, 100g/100 mL at 100°C; aqueous solution strongly acidic, pH of 0.1 M solution 1.4; insoluble in liquid ammonia; decomposed by alcohol into sodium sulfate and sulfuric acid

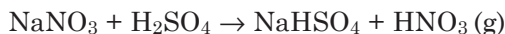
Thermochemical Properties

ΔH_f°	−269.0 kcal/mol
ΔG_f°	−237.3 kcal/mol

S° 27.0 cal/deg mol

Production

Sodium bisulfate is a by-product of sodium sulfate manufacture. One process involves reacting sulfuric acid with sodium nitrate at high temperature to form nitric acid and sodium bisulfate:



In the above reaction, nitric acid is obtained as vapor. It is purged from the system and collected in water to obtain nitric acid solution of desired concentration. Sodium bisulfate is separated by fractional crystallization.

Analysis

Elemental composition: Na 19.15%, S 26.71%, H 0.84%, O 53.30% An aqueous solution is analyzed to determine sodium content. Bisulfate anion can be measured by ion chromatography. The HSO_4^- can be measured quantitatively by titrating its aqueous solution (strongly acidic) with a standard solution of base.

SODIUM BOROHYDRIDE

[16940-66-2]

Formula NaBH_4 ; MW 37.833

Synonym: sodium tetrahydroborate

Uses

Sodium borohydride is used mostly as a reducing agent in a number of organic synthetic reactions. It reduces aldehydes, ketones and acid chlorides. The salt also is a source of hydrogen and is used to prepare other borohydrides. Other uses are bleaching wood pulp, removal of mercury from effluent wastes, decolorizing plasticizers, and as a blowing agent for plastics.

Physical Properties

White cubic crystals; hygroscopic; density 1.07 g/cm³; decomposes slowly at about 400°C in vacuum or in moist air; soluble in water, decomposing and evolving hydrogen; also soluble in alcohols, liquid ammonia, amines and pyridine.

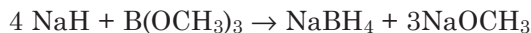
Thermochemical Properties

ΔH_f°	-45.1 kcal/mol
ΔG_f°	-29.6 kcal/mol
S°	24.2 cal/deg mol
C_p	20.7 cal/deg mol

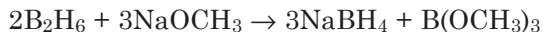
Preparation

Sodium borohydride is prepared by reacting sodium hydride with trimethyl

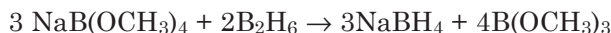
borate at about 250°C:



Also, sodium borohydride can be made by passing diborane, B_2H_6 , through a solution of sodium methylate, NaOCH_3 , in methanol:



Alternatively, diborane may be passed through a solution of sodium tetramethoxyborohydride at low temperatures:



Reactions

Sodium borohydride liberates hydrogen in contact with water, alcohol, and several other compounds. Because of its ability to release hydrogen readily, this salt is a very effective reducing agent.

Analysis

Elemental composition: Na 60.77%, B 28.58%, H 10.65%. Sodium and boron content can be measured by AA or ICP measurement. The borohydride should be dissolved cautiously in water for the metal analysis. The compound is treated with ethanol and volume of liberated hydrogen is measured to determine hydrogen content.

Hazard

Contact with oxidizers can produce violent reactions. The compound is a fire hazard because of its easy hydrogen release.

SODIUM BROMIDE

[7647-15-6]

Formula: NaBr : MW 102.89; forms a dihydrate, $\text{NaBr} \cdot 2\text{H}_2\text{O}$ [13466-08-5], MW 138.92

Occurrence and Uses

Sodium bromide occurs in seawater at an average concentration of 0.008%. It also is found naturally in some salt deposits. It is used in photography for preparing light-sensitive silver bromide emulsions. The salt also is used as a bleaching and disinfecting agent for water treatment in swimming pools, health spas, and hot tubs. Other uses are as a catalyst for partial oxidation of hydrocarbons, for increasing density of aqueous drilling fluids for oil wells, as an electrolyte component in sodium-halogen batteries, as a brominating agent in organic synthesis, in preparing bromide salts, and as a laboratory reagent. Sodium bromide is used in medicine as a sedative and hypnotic.

Physical Properties

White crystalline powder or granules; saline and slight bitter taste; cubic structure; density 3.20 g/cm³; melts at 747° C; vaporizes at 1,390°C; vapor pressure 1 torr at 806°C and 5 torr at 903°; highly soluble in methanol, 16.7 g/100mL.

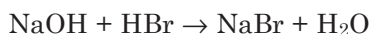
The dihydrate is a white crystalline solid; density 2.18 g/cm³; decomposes at 36°C; soluble in water; sparingly soluble in methanol.

Thermochemical Properties

$\Delta H_f^\circ(\text{cry})$	-86.30 kcal/mol
$\Delta H_f^\circ(\text{gas})$	-34.20 kcal/mol
$\Delta G_f^\circ(\text{cry})$	-83.41 kcal/mol
$\Delta G_f^\circ(\text{gas})$	-42.33 kcal/mol
$S^\circ(\text{cry})$	20.75 cal/deg mol
$S^\circ(\text{gas})$	57.65 cal/deg mol
$C_p(\text{cry})$	12.28 cal/deg mol
$C_p(\text{gas})$	8.68 cal/deg mol
ΔH_{fus}	6.24 kcal/mol

Preparation

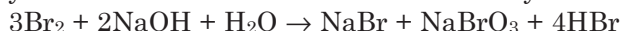
Sodium bromide can be prepared by several methods. Pure salt can be made by neutralizing sodium hydroxide or sodium carbonate with hydrobromic acid. The solution is evaporated for crystallization:



Sodium bromide can be made by passing bromine through an aqueous solution of sodium hydroxide or carbonate in the presence of a reducing agent, such as ammonia, hydrazine, activated charcoal, or Fe²⁺ ion. A typical method involves adding iron to bromine water to form ferrosferic bromide, Fe[FeBr₅]. This double salt is dissolved in excess water followed by addition of sodium carbonate. The product mixture is filtered and the filtrate is evaporated to crystallize sodium bromide. The overall reaction may be written as follows:



Another method involves adding excess bromine to a solution of sodium hydroxide. This forms sodium bromide and bromate. The product solution is evaporated to dryness. The bromate is reduced to bromide by heating with carbon:

**Analysis**

Elemental composition: Na 22.35%, Br 77.65%. The salt is dissolved in water. The aqueous solutions are analyzed for sodium by AA or ICP and for

bromide by ion chromatography. The titrimetric, colorimetric, and electrode tests for bromide ion are susceptible to interference from chloride ion. Ion chromatography should be the most reliable confirmatory test.

SODIUM CARBONATE

[497-19-8]

Formula: Na_2CO_3 ; MW 105.99; forms a monohydrate $\text{Na}_2\text{CO}_3 \cdot \text{H}_2\text{O}$ [5968-11-6], MW 124.00 and a decahydrate, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ [6132-02-1] having a molecular weight 286.14

Synonyms: The anhydrous salt Na_2CO_3 also is called "Solvay soda" and "soda ash" (technical grade is about 99% purity). The decahydrate $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$ also is known as "washing soda" or "sal soda." These two names usually refer to the technical product. Other synonyms for the decahydrate are "soda" and "Nevite."

Occurrence and Uses

Sodium carbonate occurs in nature as monohydrate in the mineral thermonatrite. It also occurs naturally as the mineral natron or natrite in its decahydrate form.

Sodium carbonate is one of the most important salts of sodium, used in manufacturing several other sodium salts. Other major uses are in manufacturing glass, soaps and detergents, pulp, and paper. Also, it is used for washing textiles and wool, in cleaning preparations, for bleaching linen and cotton, in water treatment, and in photography. Sodium carbonate is used as an emetic. Sodium carbonate solution cleanses skin and softens skin rashes. The salt is a common laboratory reagent with wide applications in analytical chemistry.

Physical Properties

The anhydrous salt is an odorless white powder; alkaline taste; hygroscopic; density 2.54 g/cm^3 ; melts at 851°C ; begins to lose CO_2 well before melting; soluble in water; insoluble in alcohol; dissolves in acids liberating CO_2 .

The monohydrate consists of colorless and odorless small crystals or crystalline powder; orthorhombic structure; refractive index 1.420; hardness 1.3 Mohs; density 2.25 g/cm^3 ; loses water at 100°C becoming anhydrous; very soluble in water; insoluble in ethanol.

The decahydrate consists of transparent crystals; effloresces on exposure to air; density 1.46 g/cm^3 ; decomposes at 34°C ; very soluble in water; insoluble in ethanol.

Aqueous solutions are strongly alkaline.

Thermochemical Properties

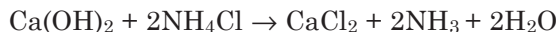
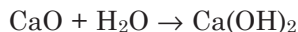
ΔH_f°	-270.2 kcal/mol
ΔG_f°	-249.6 kcal/mol

S°	32.27 cal/deg mol
C_p	26.84 cal/deg mol
ΔH_{fus}	7.10 kcal/mol

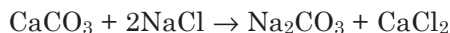
Production

Sodium carbonate at present is mostly mined from its natural deposits. It also is manufactured synthetically by Solvay (or ammonia-soda) process. The natural production of sodium carbonate currently has surpassed its synthetic production.

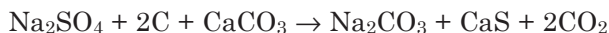
The Solvay process involves a series of partial reactions. The first step is calcination of calcium carbonate to form lime and CO_2 . Lime is converted to calcium hydroxide. The most crucial step of the process involves reacting brine solution with carbon dioxide and ammonia to produce sodium bicarbonate and ammonium chloride. Sodium bicarbonate converts to sodium carbonate. The calcium hydroxide and ammonium chloride react to form calcium chloride as the by-product. The partial reactions are shown below:



The overall reaction:



Sodium carbonate was made historically by the Leblanc process. The first commercial production was carried out by the Leblanc process. In this process, sodium chloride was treated with sulfuric acid to produce sodium sulfate and hydrochloric acid. Heating the sodium sulfate with coal and limestone produced a "black ash" that contained sodium carbonate, calcium sulfide, unreacted coal, and calcium carbonate. Sodium carbonate was separated from the black ash by leaching with water. The overall reaction is as follows:



Analysis

Elemental composition: Na 43.39%, C 11.33%, O 45.29%. Aqueous solution of sodium carbonate is strongly alkaline and its normality can be measured by acid-base titration. Sodium content can be measured by AA, ICP, and other instrumental analyses. Carbonate anion can be measured by ion chromatography or from carbon dioxide liberated when the salt is treated with dilute

acid. Liberated CO_2 can be identified by the limewater test or by GC-TCD or GC/MS (m/z 44).

SODIUM CYANIDE

[143-33-9]

Formula: NaCN ; MW 49.008

Uses

Sodium cyanide is used in extracting gold and silver from their ores. It forms soluble complexes with these metals. Other uses are in electroplating baths, heat treatment of metals, fumigation, and preparing other cyanide salts and complexes.

Physical Properties

White cubic crystals; hygroscopic; density 1.6 g/cm^3 ; melts at 563°C ; very soluble in water; aqueous solution strongly alkaline and decomposes rapidly.

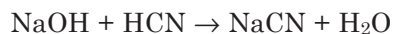
Thermochemical Properties

ΔH_f°	-20.9 kcal/mol
ΔG_f°	-18.3 kcal/mol
S°	27.6 cal/deg mol
C_p	16.8 cal/deg mol

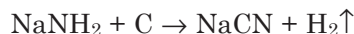
Preparation

Sodium cyanide can be prepared by several methods (See Potassium Cyanide).

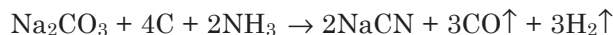
It is prepared by passing hydrogen cyanide through a 50% aqueous solution of sodium hydroxide followed by evaporation of the solution in vacuum:



Another method is to reduce sodamide with carbon at red heat:



Also, sodium cyanide can be made by heating a mixture of sodium carbonate and carbon with ammonia at high temperatures:



Reactions

Reactions of sodium cyanide are similar to those of potassium cyanide (See Potassium Cyanide).

Analysis

Elemental composition: Na 46.92%, C 24.50%, N 28.58%. An aqueous solution is analyzed for sodium. Cyanide is measured by an electrode specific to cyanide ion. Alternatively, cyanide may be measured by pyridine-barbituric acid colorimetric test (See Hydrogen Cyanide).

Toxicity

Sodium cyanide is extremely toxic. Ingestion of a small quantity can be fatal. The toxic properties are similar to Potassium Cyanide (See Potassium Cyanide).

SODIUM ETHOXIDE

[141-52-6]

Formula: $\text{C}_2\text{H}_5\text{ONa}$; MW 68.06

Synonyms: sodium ethylate; caustic alcohol

Uses

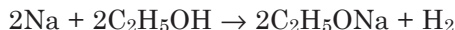
Sodium ethoxide is used in organic synthesis for condensation reactions. It also is a catalyst in many organic reactions.

Physical Properties

White or yellowish powder; hygroscopic; darkens and decomposes on exposure to air; decomposes in water forming sodium hydroxide and ethanol; dissolves in absolute ethanol.

Preparation

Sodium ethoxide is prepared by reacting sodium with absolute ethanol:



Sodium in small quantities is added to absolute alcohol at 10°C. The temperature is raised to warming (to about 38°C). The mixture is cooled again and sodium and absolute alcohol are added gradually followed by careful warming. The process is repeated to obtain a sufficient yield of the product.

Analysis

Elemental composition: Na 33.79%, C 35.29%, H 7.41%, O 23.51%. The compound is decomposed in water cautiously. A portion of the aqueous solution is measured for sodium hydroxide by acid-base titration, while another portion is analyzed for sodium by AA or ICP.

SODIUM FLUORIDE

[7681-49-4]

Formula: NaF; MW 41.988

Uses

Sodium fluoride is used in electroplating, as a steel degassing agent, in vitreous glasses and enamels, in heat-treating salt compositions, and preserving wood. The salt also is used in pesticide formulations and as an insecticide for ant and roach control. Sodium fluoride is used for fluoridating drinking water and for disinfecting apparatus in distilleries. An important application of this salt is preparing other fluoride salts. Sodium fluoride occurs in nature as the mineral villiaumite.

Physical Properties

Colorless cubic or tetragonal crystals; density 2.78 g/cm³; melts at 993°C; vaporizes at 1,695°C; moderately soluble in water 4.22 g/100mL at 18°C; soluble in hydrofluoric acid; insoluble in ethanol.

Thermochemical Properties

ΔH_f°	-137.1 kcal/mol
ΔG_f°	-129.9 kcal/mol
S°	12.3 cal/deg mol
C_p	11.2 cal/deg mol

Preparation

Sodium fluoride is prepared by adding sodium hydroxide or sodium carbonate to a 40% solution of hydrofluoric acid. In excess hydrofluoric acid, sodium bifluoride, NaHF₂, is formed. NaF also is made by fusion of cryolite with caustic soda. Technical grade products are usually sold at 90 to 95% purity.

Analysis

Elemental composition Na 54.75%, F 45.25%. The salt is dissolved in water and analyzed for sodium and fluoride anion. The anion can be measured effectively with a fluoride ion-selective electrode or by ion chromatography.

Toxicity

Sodium fluoride is an acute toxicant. Ingestion of large quantities (5 to 10g) can cause death in humans. Smaller quantities can produce nausea, vomiting, diarrhea, stupor, and weakness. Other symptoms are tremor, muscular weakness, and dyspnea. Mottling of teeth can occur from chronic exposure.